

## How to Make a Buffer

Making Your Own Buffer, Part I:

- The most effective buffer contains large, equal amounts of the proton donor and proton acceptor.
- The capacity of a buffer (its ability to continue to absorb any acid or base added) is dependent upon the concentrations of the two components of the buffer; the higher the concentrations, the higher the capacity.
- If you have equal amounts of the weak acid and conjugate base:

a.  $[HA] = [A^-]$

b. Therefore, your  $K_a$  equation would simplify to:

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = [H_3O^+]$$

c. And thus, **to get the best possible buffer, you want to choose a weak acid such that**

$$\frac{pK_a}{\text{of weak acid}} = \frac{pH}{\text{of buffer sol'n}} \quad (\text{because } K_a = [H^+])$$

**Example:** An industrial process requires a constant pH of 3.00. The weak acids available in the warehouse are:

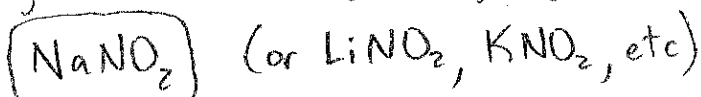
benzoic acid ( $HC_7H_3O_2$ ), $K_a = 6.3 \times 10^{-5}$	$\frac{pK_a}{4.20}$
nitrous acid ( $HNO_2$ ), $K_a = 7.2 \times 10^{-4}$	$3.14$
phosphorous acid ( $H_3PO_3$ ), $K_a = 3.7 \times 10^{-2}$	$1.43$
hydrocyanic acid ( $HCN$ ), $K_a = 6.2 \times 10^{-10}$	$9.21$

a. Which of the four acids would be the best choice for the industrial process? Explain with a calculation.

$HNO_2$ , b/c when  $[HNO_2] = [NO_2^-]$ , the pH of the buffer sol'n created would be 3.14, which is closer to the desired pH of 3.00 than the other options.

b. What other chemical(s) are needed to prepare the buffer? Give at least one possible chemical which could be used.

You need the conjugate base of  $HNO_2$ ,  $NO_2^-$ , so you could add:



c. How much of the chemical from (b) should you add to achieve a buffer with a pH of 3.00?

$$pH = 3.00 \Rightarrow [H_3O^+] = 1.0 \times 10^{-3} M$$

You need  $\frac{0.72 \text{ mol } NaNO_2}{1.0 \text{ mol } HNO_2}$  to get pH = 3.00

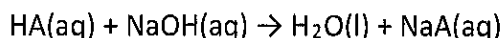
$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} \Rightarrow \frac{[NO_2^-]}{[HNO_2]} = \frac{K_a}{[H_3O^+]} = \frac{7.2 \times 10^{-4}}{1.0 \times 10^{-3}} = 0.72$$

1. A student wishes to create a buffered solution with a pH of 5. Which of the following acids would be the best choice for the buffer?

- $H_2C_2O_4$   $K_a = 5.9 \times 10^{-2}$
- $H_3AsO_4$   $K_a = 5.6 \times 10^{-3}$
- $HC_2H_3O_2$   $K_a = 1.8 \times 10^{-5}$
- $HOCl$   $K_a = 3.0 \times 10^{-8}$

## Making Your Own Buffer, Part II:

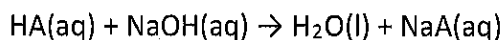
Let's see what happens if you add a strong base, NaOH, to the weak acid, HA:



Welp, it got neutralized. (I know, I know: what else did I expect?)

But WAIT! Let's see what happens if I add only 1/2 the amount of strong base:

**Now only *half* of the weak acid is neutralized:**



**but the other half of the weak acid, HA, is still there!**

Did you see what just happened?? After adding half the amount of strong base to a given amount of weak acid, the species floating around in my reaction vessel include:

HA, NaA, and H<sub>2</sub>O

## That's a buffer!!!!

This means that any titration involving a weak acid and/or weak base is a buffer problem.

→ If I titrate HA with NaOH, then as soon as the first drop splashes into the container, I've made NaA which promptly dissociates into A<sup>-</sup>. More on titrations soon!

Now we have lots of different ways to create a buffer! But be careful: if you're mixing a weak acid/base with a strong base/acid to make a buffer, you have to be careful to not to completely neutralize your weak species.

	Species 1	Species 2	Buffer?
Example 1	100 mL 0.20 M HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	100 mL 0.10 M NaOH	<b>Yes</b> , moles of strong base < moles of weak acid.
Example 2	100 mL 0.20 M HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	50 mL 0.20 M NaOH	<b>Yes</b> , moles of strong base < moles of weak acid.
Example 3	100 mL 0.20 M HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>	100 mL 0.20 M NaOH	<b>Nope</b> , moles of strong base completely neutralized weak acid.
Example 4	100 mL 0.20 M HNO <sub>3</sub>	100 mL 0.10 M NaOH	<b>Nope</b> , strong acids and bases don't make buffers!

Circle all the combinations below that would make a buffer solution when mixed. Note: in most cases, a neutralization reaction will occur. Consider what the products of the neutralization will be and how much of each species will be present when the reaction is complete.

X 100 mL of 1.0 M HCl with 100 mL of 1.0 M NaOH  
*SA + SB*

X 100 mL of 1.0 M HCl with 50 mL of 1.0 M NaOH  
*SA + SB*

X 100 mL of 1.0 M HNO<sub>2</sub> with 100 mL of 1.0 M NaOH  
*WA = SB*

100 mL of 1.0 M HNO<sub>2</sub> with 50 mL of 1.0 M NaOH  
*WA > SB*

100 mL of 1.0 M NaC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> with 100 mL of 0.50 M HCl  
*WB > SA*

X 50 mL of 1.0 M HC<sub>2</sub>H<sub>3</sub>O<sub>2</sub> with 100 mL of 0.50 M LiOH  
*WA = SB*

$$4 < pK_a < 5$$

## Free Response Practice

1. A chemist creates a buffer at pH 4.30 by combining one of the acids from the following list and their soluble salt in a solution.

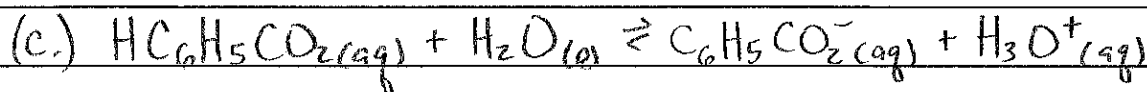
		$K_a$	$pK_a$
• chloroacetic acid, $HC_2H_2ClO_2$	$K_a = 1.35 \times 10^{-3}$		
• propanoic acid, $HC_3H_5O_2$	$K_a = 7.3 \times 10^{-4}$		3.13
• benzoic acid, $HC_6H_5CO_2$	$K_a = 6.4 \times 10^{-5}$		4.19
• hypochlorous acid, $HClO$	$K_a = 3.5 \times 10^{-8}$		

- Which weak acid would work best? Justify your choice with calculations.
- Identify an appropriate salt that can be used with the weak acid chosen above to complete the buffer system.
- Write an equilibrium reaction to describe this buffer.
- What would be true about the weak acid/conjugate base ratio ( $\frac{HA}{A^-}$ ) in your buffer system if the acid chosen above was used to prepare the desired buffer at pH 4.30? Justify your answer.
- How would adding HBr to the buffer system change the weak acid/conjugate base ratio ( $\frac{HA}{A^-}$ ) identified in part (d)? Explain.
- How could you prepare a buffer solution from the weak acid chosen in part (a) if you were NOT provided with the conjugate salt, but instead solutions of a strong acid and a strong base?

(a.) Benzoic acid, b/c when  $[HC_6H_5CO_2] = [C_6H_5CO_2^-]$ , the

pH of the buffer sol'n created would be 4.19, closer to the desired pH of 4.30 than the other options.

(b.)  $NaC_6H_5CO_2$



(d.) Since the desired pH of 4.30 > 4.19, the  $pK_a$  of the weak acid (and thus the pH of sol'n if  $[HC_6H_5CO_2] = [C_6H_5CO_2^-]$ ),

$\frac{HA}{A^-} < 1$ , since the sol'n is more basic than @ pH =  $pK_a$ , so  $[A^-] > [HA]$ .

(e.) Adding HBr is adding  $H^+$  ions, which will convert the conjugate base ( $A^-$ ) into weak acid (HA),  $\uparrow [HA]$  and  $\downarrow [A^-]$ . This would increase the HA/ $A^-$  ratio, possibly to 1 or > 1 depending on how much HBr was added.

(f.) Add strong base to the weak acid until # mol SB =  $\frac{1}{2}$  # mol WA (ignore the strong acid  $\rightarrow$  not helpful)